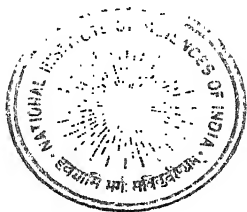


NISI MONOGRAPHS

**CHEMISTRY  
OF  
BI- AND TRIPOSITIVE SILVER**

BY  
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AND  
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**SILVER JUBILEE ISSUE**

1960

**National Institute of Sciences of India**





## PREFACE

Silver in its higher oxidation states may be said to have been the subject of study from more than a century and half ago, though no serious attention was paid to the problem till towards the close of the nineteenth century. Since then, however, there has accumulated a more or less voluminous literature on the subject of a rather limited character and some facts of highly interesting and exciting nature have been revealed now and then from researches in the field of coordination complexes. Barring one or two short reviews there has been no attempt to furnish a systematic and comprehensive account of the subject till now. That such an account will be of considerable use to the workers in this line needs little emphasis to state here, as the results already achieved promise a richer harvest in future.

We have, therefore, taken advantage of this occasion of celebrating the Silver Jubilee of the National Institute of Sciences of India to prepare the present monograph on the subject with the hope that it might fulfil the need felt by those engaged in researches on the chemistry of silver in its unusual valency states.

Our best thanks are due to Dr. Supriya Roy for the microphotographs included in this work.

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## ERRATA

1. In line 4 of the Preface after *century for* , *read* .
2. P. 1, paragraph 2 in line 3, for  $(n-1)d^{10}s^2$  *read*  $(n-1)d^{10}ns^1$ .  
„ paragraph 2 in line 4, *for* 3, 4 and 5 *read* 4, 5 and 6.
3. P. 5, in line 3 from the bottom for  $(x_1-x_2)$  *read*  $(x_2-x_1)$ .
4. P. 22, in line 9 *for* oxidiiing *read* oxidizing.
5. P. 32, reference No. 21, for Noyts *read* Noyes.





# CONTENTS

	PAGE
PREFACE                    ...            ...            ...	iii
INTRODUCTION            ...            ...            ...	I
I. SIMPLE COMPOUNDS    ...            ...            ...	
( <i>a</i> ) Oxides of Silver(II) and Silver(III) ; Silver Peroxysalts    ...            ...	3
( <i>b</i> ) Silver(II) Fluoride    ...            ...	12
II. COMPLEX COMPOUNDS    ...            ...	13
( <i>a</i> ) Silver(II) Complexes                    ...	14
( <i>b</i> ) Silver(III) Complexes                    ...	23
REFERENCES                ...            ...            ...	30



# CHEMISTRY

## OF

### BI- AND TRIPOSITIVE SILVER

#### INTRODUCTION

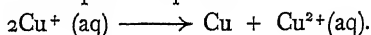
Silver, an element of the second transition series, shows normally a positive valency of 1 in keeping with its position in Group I, Subgroup B, of Mendeleef's Periodic Table. But there are compounds where it is known to occur with a higher positive valency of 2, which is the common oxidation state of its lower homologue copper in the same subgroup. Compounds are also known in which silver manifests a still higher valency of 3, so characteristic of its higher homologue gold. These higher oxidation states of silver are particularly stabilized, as in the case of many other elements, by complex formation; for, the tendency of  $\text{Ag}^{2+}$  and  $\text{Ag}^{3+}$  to resume the normal state of  $\text{Ag}^+$  by capture of one or two electrons is opposed by the screening effect of an envelope of negative charges from the polar coordinating ligands.

The chemical analogy of silver with copper and gold in the same subgroup rests obviously on their having the same configuration of the outermost electrons,  $(n-1)d^{10}s^2$ , and the same  $^2\text{S}$  ground state. The value of  $n$  for copper, silver and gold is given respectively by 3, 4 and 5. A comparison of their ionization and electrode potentials reveals the relative stability of their ions in the different states of valency as shown below.

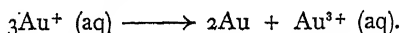
	Cu	Ag	Au
Ionization potential $I(1) - (\text{eV})$	7.72	7.87	9.22
$E^\circ \text{ M/M}^+ (\text{V})$	0.522	0.799	1.68
$\text{M/M}^{2+}$ „	0.345	1.390	
$\text{M/M}^{3+}$ „			1.42
$\text{M}^+/\text{M}^{2+}$ „	-1.70	2.0	

The electrode potential values indicate that the silver ions should become increasingly unstable with increasing charge, showing an increasing tendency for reduction, and that the reverse would hold good for copper and gold. This is supported by observed facts. For, while unipositive silver compounds are quite stable and compounds of bi- and tripositive silver are easily reduced, those of bipositive copper and tripositive gold are much more stable than their corresponding unipositive compounds. This is accounted for by the

fact that the heat of hydration for  $\text{Cu}^{2+}$  compared with that for  $\text{Cu}^+$  makes the cupric compounds in aqueous solution more stable, in spite of the quite high second ionization energy (20.2 eV) of copper. It is well known that cuprous compounds in aqueous solution suffer an easy disproportionation into metallic copper and cupric compounds:



This differs strikingly from the behaviour of  $\text{Ag}^+$  ion. Unipositive gold also behaves like copper in this respect:



A comparison of the equilibrium constant of the cupric (II)—cuprous(I) system with that of the argentic (II)—argentous (I) system reveals this difference quite clearly.

$$\frac{[\text{Cu}^{2+}]}{[\text{Cu}^+]^2} = 1.2 \times 10^6; \quad \frac{[\text{Ag}^{2+}]}{[\text{Ag}^+]^2} = 10^{-20} \text{ at } 25^\circ\text{C}.$$

In complex formation both copper and silver can occur in the tripositive state; but curiously enough the compounds of tripositive silver are more stable than those of tripositive copper, though there is not much difference between the two elements in their second and third ionization potentials. Thus while in the case of copper the stability of its valency states follows the order bi- > uni- > trivalent, that in the case of silver is given by uni- > bi- and trivalent. Then again for the same valency state the stability order for the two elements can be represented as  $\text{Ag(I)} > \text{Cu(I)}$ ,  $\text{Cu(II)} > \text{Ag(II)}$ , and  $\text{Ag(III)} > \text{Cu(III)}$ . Another interesting observation made in this connection is the phenomenon of direct oxidation of  $\text{Ag(I)}$  to  $\text{Ag(III)}$  in many cases without the intermediate formation of  $\text{Ag(II)}$ , which is then formed from  $\text{Ag(III)}$  by subsequent reduction (Noyes *et al.* 1935). The disproportionation of  $\text{Ag(II)}$  into  $\text{Ag(III)}$  and  $\text{Ag(I)}$ , demonstrated by some workers (Barbieri 1931; Gordon and Wahl 1958), is presumably related to it. In the case of gold the bipositive state is practically unknown. The difference, thus exhibited by silver in its different oxidation states from its lower and higher homologues copper and gold in spite of similar electronic configurations (viz.  $d^{10}$  for uni-,  $d^9$  for bi- and  $d^8$  for tripositive state), is rather intriguing.

In the present work we propose, therefore, to make a comprehensive and critical treatment of the chemistry of bi- and tripositive silver in which we have been interested for several years.

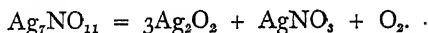
Reference may also be made here to a short but excellent review on the subject by Bailar (1944) that appeared in the Journal of Chemical Education nearly 16 year ago. A short, but very informative, chapter on the subject has also been included by Kleinberg (1950) in his book on the "Unfamiliar Oxidation States and their Stabilization".

## I. SIMPLE COMPOUNDS

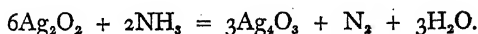
## (a) OXIDES OF SILVER(II) AND SILVER(III); SILVER PEROXYSALTS.

As early as in 1804 Ritter obtained a black crystalline powder that separated at the anode when an electric current was passed between platinum electrodes through an aqueous solution of silver nitrate. The substance was for a long time considered to be a peroxide of silver, as it was found to give off oxygen when heated. Subsequent analyses of the product by various workers (Mulder and Heringa 1896; Mulder 1897, 1898, 1903; Sulc 1896, 1900; Tanatar 1901; Watson 1906; Barbieri 1906, 1907) almost after a century revealed the presence of nitrogen in the compound with a composition represented by the formula,  $\text{Ag}_7\text{NO}_{11}$  or  $3\text{Ag}_2\text{O}_2 \cdot \text{AgNO}_3 \cdot 2\text{O}$ . The name silver peroxyxynitrate was generally given to it. This was formulated as  $2\text{Ag}_3\text{O}_4 \cdot \text{AgNO}_3$  by Tanatar (1901). By the electrolysis of solutions of other salts of silver, or of acids with a silver anode, peroxy sulphate of the composition,  $5\text{Ag}_2\text{O}_2 \cdot 2\text{Ag}_2\text{SO}_7$  (Mulder 1900), and the peroxyfluorides,  $\text{Ag}_{15}\text{F}_3\text{O}_{16}$  or  $4\text{Ag}_3\text{O}_4 \cdot 3\text{AgF}$  and  $\text{Ag}_7\text{FO}_8$  or  $2\text{Ag}_3\text{O}_4 \cdot \text{AgF}$  (Tanatar 1901), were obtained. The black crust, which was observed by Schiel (1864) to separate on a silver anode during the electrolysis of dilute sulphuric acid, was very probably this peroxy sulphate and not silver peroxide or dioxide.

Wöhler (1867) had described the preparation of a silver dioxide by the action of ozone on silver monoxide or silver, and an oxide of the composition,  $\text{Ag}_4\text{O}_8$ , was believed by Berthelot (1880) to result from the action of alkaline  $\text{H}_2\text{O}_2$  on silver nitrate solution. Marshall (1900), as also Marshall and Inglis (1902), reported the formation of silver peroxide as a black precipitate by the action of potassium persulphate on silver nitrate or silver sulphate solution. The decomposition of silver peroxyxynitrate with boiling water was found to give oxygen and silver nitrate with separation of silver peroxide (Mulder 1898; Watson 1906) according to the equation:



The peroxide,  $\text{Ag}_2\text{O}_2$ , reacts with ammonia, oxidizing the latter to nitrogen (Watson 1906):

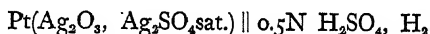


The oxide,  $\text{Ag}_4\text{O}_8$ , has, however, not been isolated.

That these oxidation products of silver are not true peroxides, but a compound or a mixture of compounds of silver in a higher oxidation state, was demonstrated by Barbieri (1906, 1907). For, an acid solution of these oxidation products was found by him not to be affected by permanganate, manganese dioxide or lead peroxide, contrary to peroxy compounds. Moreover, they were not found to give any hydrogen peroxide on treatment with acids as could be proved by titanium test (Yost 1926). The black precipitate

formed by the action of potassium persulphate on silver nitrate solution was, however, found to consist of a mixture of  $\text{Ag}_2\text{S}_2\text{O}_8$  and  $\text{Ag}_2\text{O}_2$  (Austin 1911).

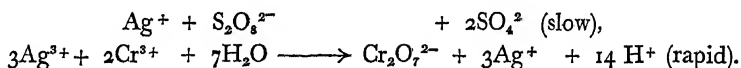
Luther and Pokorny (1908) have shown that in alkaline solution silver or  $\text{Ag}^+$  is reversibly and quantitatively converted by anodic oxidation only to the bipositive,  $\text{Ag}^{2+}$ , state, while in acid solution the oxidation goes up to the tripositive,  $\text{Ag}^{3+}$ , state. They measured the oxidation potentials of the systems  $\text{AgO}-\text{Ag}_2\text{O}$ , and of  $\text{Ag}_2\text{O}_3-\text{AgO}$ , as well as of  $\text{Ag}_2\text{O}_3-\text{Ag}_2\text{SO}_4$  (saturated). Their results have been confirmed by many subsequent workers (Baborovsky and Kuzma 1909; Brown 1916; Jirsa *et al.* 1926; Barbieri 1931; Carman 1934). Silver peroxyxynitrate, when reduced in 1N NaOH, shows an indication of the existence of tripositive silver. Luther and Pokorny (1908) observed three different potentials of 1.57 volts, 1.41 volts and 1.17 volts. The last two corresponded to the known values for the change of  $\text{AgO} \rightarrow \text{Ag}_2\text{O}$  and  $\text{Ag}_2\text{O} \rightarrow \text{Ag}$  states respectively. The higher value of 1.57 volts was, therefore, considered to correspond to a reduction from a still higher state, and in all probability from  $\text{Ag}_2\text{O}_3$  to  $\text{AgO}$ . They have further found that in acid solution the compound  $\text{Ag}_2\text{O}_3$  is directly decomposed into  $\text{Ag}$  on electrolytic reduction and conversely  $\text{Ag}$  can be oxidized directly to  $\text{Ag}_2\text{O}_3$ . A potential of 1.74 volt has been recorded by them for the E.M.F. of the cell



From a determination of the transference number of silver in a solution of  $\text{AgO}$  in nitric acid Weber (1917) concluded that at least a portion of the oxide should be considered as an oxide of tripositive silver.

The formation of silver(III) by the action of dry ozone on dry pulverized silver was established from potential measurement by Jirsa *et al.* (1926). The product thus obtained was not pure  $\text{Ag}_2\text{O}_3$ , but a solid solution of  $\text{Ag}_2\text{O}_3$ ,  $\text{Ag}_2\text{O}_2$  and  $\text{Ag}_2\text{O}$ , their relative proportions depending on the condition for oxidation.

The marked catalytic effect of  $\text{Ag}^+$  ion in certain oxidations by persulphate is attributed to the formation of an unstable oxide,  $\text{Ag}_3\text{O}_4$  (Higson 1921), or of the unstable  $\text{Ag}^{3+}$  ion (Yost 1926; Yost and Claussen 1931; Bekier and Kijowski 1934; Dekker, Levy and Yost 1937). The mechanism of this catalytic oxidation, for instance in the case of  $\text{Cr(III)}$  to  $\text{Cr(VI)}$ , is represented as follows:



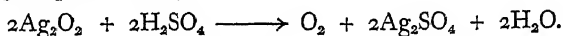
Malaguti (1952), on the other hand, attributes this catalytic effect to the  $\text{Ag}^{2+}$  ion in contradiction to Yost's view. He has shown that *bis-aa'*-dipyridyl silver(I) ion, which is oxidized by persulphate to the corresponding complex

of silver(II) ion, is as effective a catalyst as  $\text{Ag}^+$  ion. A similar view has been advanced by Bawn and Margerison (1955). It might, however, be argued that it is not  $\text{Ag}^{2+}$  ion but  $\text{Ag}^{3+}$  ion, arising out of dismutation of  $\text{Ag}^{2+}$  into  $\text{Ag}^+$  and  $\text{Ag}^{3+}$ , which actually acts as the catalyst here (cf. Barbieri 1931).

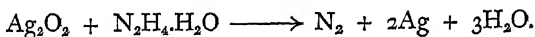
The black precipitate formed by the action of persulphate on silver nitrate solution, when rapidly analysed, shows the presence of silver in the tripositive state in the product (Yost 1926; Carman 1934), and the substance was believed to be either  $\text{Ag}_2\text{O}_3$  or a basic  $\text{Ag(III)}$  sulphate (Yost 1925). On standing, it loses oxygen and changes ultimately to a product containing  $\text{Ag(III)}$  and  $\text{Ag(II)}$  in equal amounts (Carman 1934).

Many chemical and physical properties of silver(II) oxide were studied by Jirsa (1925).

On treatment with sulphuric or perchloric acid the compound dissolves at first to form a brown solution, but decomposes readily to liberate all its active oxygen quantitatively:

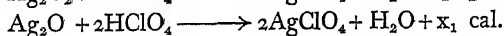


It is a powerful oxidizing agent which can liberate nitrogen from ammonia solution. Hydrazine hydrate is quantitatively oxidized by it with evolution of nitrogen and formation of metallic silver:



When dissolved in nitric acid silver(II) oxide yields a brown solution of high oxidizing power with partial decomposition into silver(I) and with liberation of oxygen. The extent of decomposition increases with temperature and dilution of the acid. From the brown solution in strong nitric acid it can be reprecipitated by dilution with water (Jirsa 1925). Similar observations were made also by Carman (1934).

The same author (Jirsa 1925) determined the heat of reaction of silver (II) oxide with perchloric acid, hydrazine hydrate and nitric acid respectively by calorimetric measurement. Similar measurements were made for the oxide,  $\text{Ag}_2\text{O}$ , as well. As the reaction of the two oxides with these reagents in dilute solution is rather slow, the measurements were made in their concentrated solutions and the values obtained were extrapolated to zero dilution in each case. The difference between the zero-dilution values of the two oxides in each case was found to be 1,000 cal. nearly.



$$(x_1 - x_2) = (10,964 - 9,964) = 1,000 \text{ cal. approx.}$$

This gives the heat of formation of  $\text{Ag}_2\text{O}_2$  from  $\text{Ag}_2\text{O}$ :

$\text{Ag}_2\text{O}_2 \longrightarrow \text{Ag}_2\text{O} + \frac{1}{2}\text{O}_2 + 1,000 \text{ cal.}$  Taking the value of the heat of

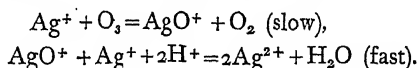
formation of  $\text{Ag}_2\text{O}$  as 6,440 cal. (Lewis) the heat of formation of  $\text{Ag}_2\text{O}_2$  from silver and oxygen was given by the author (Jirsa 1925) as 5,440 cal.; i.e.  $\text{Ag}_2\text{O}_2 \longrightarrow 2\text{Ag} + \text{O}_2 - 5,440$  cal. The necessity of extrapolation, however, introduces some amount of uncertainty in the heat values calculated. From these values of heat of formation of  $\text{Ag}_2\text{O}_2$  from  $\text{Ag}_2\text{O}$  and from  $\text{Ag}$  the author calculated by means of approx. Nernst formula the dissociation pressures of  $\text{Ag}_2\text{O}_2$  into  $\text{Ag}_2\text{O}$  and into  $\text{Ag}$  as  $3.96 \times 10^8$  and  $1.37 \times 10^3$  atm. respectively at  $25^\circ\text{C}.$

$$\log p_{\text{O}_2} = -Q/4.571 T + 1.75 \log T + 2.8.$$

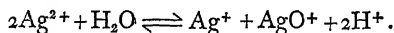
In spite of these high values of dissociation pressure the existence of  $\text{Ag}_2\text{O}_2$  in the thermodynamic sense at room temperature in a fairly stable condition must be attributed to some passivity phenomenon (Jirsa 1925).

The physicochemical investigations, however, demonstrate the existence of bipoisitive silver more precisely than the tripositive one. Luther and Pokorny (1908) have shown that the voltage-current curve for a silver anode in sodium hydroxide solution shows that the silver is quantitatively oxidized first to  $\text{Ag}_2\text{O}$  and then to  $\text{AgO}$ . Under these experimental conditions there is no indication of the formation of tripositive silver ion.

The oxidation of silver in nitric acid solution with ozone was extensively studied by Noyes *et al.* (1935). The oxidation leading to the formation of a black soluble argentic nitrate was found to reach an incomplete but steady state due to the reduction of the product by water with evolution of oxygen. From a study of the rate of oxidation at  $0^\circ\text{C}$  the authors assumed the intermediate formation of an oxygenated tripositive silver ion:



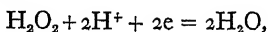
By oxidimetric titration and magnetic measurements they, however, proved that in the steady state the silver was present completely in the bipoisitive form in solution. To account for the precipitation of basic salts of silver (III) from the latter on dilution the authors proposed the following equilibrium reaction:



The standard oxidation potential for the system  $\text{Ag}^{2+} - \text{Ag}^+$  was determined by Noyes and Kossiakoff (1935) from the measurement of E.M.F. between platinum electrodes of silver(I) and silver(II) nitrates in 1-2N  $\text{HNO}_3$  against hydrogen electrode. The value was found to be  $1.914 \pm 0.002$  volts. This is, however, much higher than 1.41 volts reported earlier by Luther and Pokorny (1908), as well as by others, from determination in alkaline solution. This higher value of 1.914 volts for the standard oxidation potential of the system  $\text{Ag}^{2+} - \text{Ag}^+$  in acid solution is regarded as a support to the hypothesis

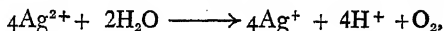


that the decomposition of bipoisitive silver by water with evolution of oxygen proceeds through the intermediate formation of hydrogen peroxide ; because the potential for the reaction,



has a value of only 1.77 volts.

The rate of reduction by water of silver(II) nitrate in nitric acid solution of various acid concentrations has been studied at 0° and 25°C by Noyes, Coryell *et al.* (1937). The rate of reduction for the overall reaction,



is represented by the equation,

$$\frac{-d(\text{Ag}^{2+})}{dt} = k_2 \frac{(\text{Ag}^{2+})^2}{(\text{Ag}^+)} + k_4 \frac{(\text{Ag}^{2+})^4}{(\text{Ag}^+)^3},$$

where both  $k_2$  and  $k_4$  decreased greatly with increase in the concentration of nitric acid. The equation was derived by assuming the occurrence of two independent reactions, both involving the presence of tripositive silver in equilibrium with uni- and bipoisitive species.

The standard oxidation potential in 4*N*  $\text{HClO}_4$  at 25°C for the reaction,  $\text{Ag}^{2+} + e = \text{Ag}^+$ , was found to have a value of 2.0 volts (Noyes, Devault *et al.* 1937). This is about 66.71 millivolts higher than that found in 4*N*  $\text{HNO}_3$  solution. While the silver(II) nitrate solution in nitric acid was found to be coloured black to brown, silver(II) perchlorate in perchloric acid solution (60 per cent or below) is, on the other hand, pink. The above observations have been explained by the assumption that in nitric acid solution the bipoisitive silver mainly exists in the form of a nitrate complex, because the solubility of  $\text{Ag}(\text{II})$  increases with the concentration of  $\text{NO}_3^-$  ion at a constant hydrogen ion concentration. The passage of some silver to the anode during transference experiment in the dark brown solutions obtained by dissolving silver(II) oxynitrate in concentrated nitric acid is also in support of the above assumption that bipoisitive silver in concentrated nitric acid exists mainly in the form of a nitrate complex (Weber 1917). Silver(II) oxide,  $\text{AgO}$ , is converted in 3*N*  $\text{HNO}_3$  into silver(II) oxynitrate after equilibrium is attained. The empirical composition of the product corresponded to the formula,  $\text{AgO}_{1.48} (\text{NO}_3)_{0.153}$ . A more or less similar product was obtained by electrolysis by earlier workers.

By the electrolysis of  $\text{AgClO}_4$  in perchloric acid solution the precipitate of an unstable oxide or a basic salt of tripositive silver was said to be produced. This substance appeared also to be formed on the surface of  $\text{AgO}$ , when the latter was treated with perchloric acid (Noyes, Devault *et al.* 1937).

From the standard oxidation potentials measured at 0°C and 25°C,

Noyes, Devault *et al.* (1937) calculated the thermodynamic constants for the reactions in the cells:

- (i)  $\text{Pt} (\text{H}_2, \text{H}^+) \parallel (\text{Ag}^{2+}, \text{Ag}^+ (4M \text{HNO}_3))$   
 (ii)  $\text{Pt} (\text{H}_2, \text{H}^+) \parallel (\text{Ag}^{2+}, \text{Ag}^+ (4M \text{HClO}_4))$

The values obtained are: (i)  $\Delta F^\circ = -44.49 \text{ kcal.}$ ;  $\Delta H^\circ = -41.0 \text{ kcal.}$   
 (ii)  $\Delta F^\circ = -46.13 \text{ kcal.}$ ;  $\Delta H^\circ = -41.1 \text{ kcal.}$

Silver (II) oxide,  $\text{AgO}$ , has also been prepared by other methods besides the hydrolysis of silver peroxynitrate. Lustrous black crystals of the compound were obtained by passing fluorine into a cold solution of  $\text{AgNO}_3$  or  $\text{AgClO}_4$  or of  $\text{AgF}$  in a platinum tube (Fichter and Goldach 1930). Jirsa (1947, 1949) prepared it by adding an alkaline solution of potassium persulphate to a boiling solution of silver nitrate (cf. DeBoer and Ormond 1952).

Sen (1948) has shown that values for certain physical constants of silver like entropy, vibration frequency, the distance of closest approach of atoms, and others calculated on the basis of the bivalency of silver are in good agreement with their observed values.

Kappana and Talaty (1951) have shown that in the electrolysis of silver fluoride solution the decomposition potential curves ran parallel to the current-anode potential curve, each exhibiting two breaks; while the current-cathode potential curve gave only one break. The second break in the decomposition potential curves is said to be associated with the anodic deposition of a black crystalline compound as reported by Tanatar (1901). But they found no fluorine in the product and represented it, on the basis of its silver content, as an oxide of tetrapositive silver having the composition  $\text{AgO}_2$ . There is, however, no justification for this assumption in view of the absence of any analytical result for the oxidation state of silver or for the peroxidic oxygen in the compound. It seems quite probable that the product is a mere hydrated higher oxide of silver,  $\text{Ag}_3\text{O}_4 \cdot 2\text{H}_2\text{O}$ , as the authors did not take proper care to dry the product before analysis. The determination of the silver content of the substance was based on the analysis of moist specimens only. The substance was also found to be feebly paramagnetic with a moment value of 0.37 B.M. only.

The anodic behaviour of silver in  $2\text{NH}_2\text{SO}_4$  at  $25^\circ\text{C}$  was studied by Jones and Thirsk (1954). The oxidation was found to occur as  $\text{Ag} \longrightarrow \text{Ag}_2\text{SO}_4 \longrightarrow \text{AgO} + \text{O}_2(\text{g})$ , and the reduction of the oxidized anode to follow the sequence  $\text{AgO} \longrightarrow \text{Ag}_2\text{SO}_4 \longrightarrow \text{Ag} + \text{H}_2(\text{g})$ . An X-ray analysis of the oxide formed at the anode demonstrated its identity with specimens of  $\text{AgO}$  prepared by chemical methods, though admixed with traces of its decomposition products, namely the monoxide,  $\text{Ag}_2\text{O}$ , and a suboxide of silver.

Preparation of  $\text{AgO}$  from silver nitrate solution by controlled oxidation with alkaline hypochlorite at  $75^\circ\text{--}85^\circ\text{C}$  has been reported by Dutta (1955).

Single crystals and powdered samples of  $\text{Ag}_7\text{NO}_{11}$ , erroneously described as  $\text{Ag}_2\text{O}_3$ , have also been studied by X-ray diffraction, and a face-centred cubic structure with a cube edge of 9.870 A.U. has been derived (Brackken 1935; Barbieri and Malaguti 1950).

Neiding and Kazarnovskii (1951) studied the properties of  $\text{AgO}$ , prepared by the decomposition of the peroxyxynitrate,  $\text{Ag}_7\text{NO}_{11}$ , with boiling water. The latter was obtained by the anodic oxidation of silver nitrate solution. They observed that dilute acids ( $\text{H}_2\text{SO}_4$ ,  $\text{HNO}_3$ ) gave no  $\text{H}_2\text{O}_2$  by their action on  $\text{AgO}$ ; concentrated nitric acid gave a dark brown solution, which on dilution with water and heating evolved oxygen and became colourless. The electrical conductivity of  $\text{AgO}$ , pressed under 12,000 kg/sq. cm at  $20^\circ\text{C}$ , was found to be  $7 \times 10^{-2} \text{ ohm}^{-1} \text{ cm}^{-1} \pm 15$  per cent., the temperature coefficient between  $-40^\circ$  and  $20^\circ\text{C}$  being positive. The substance thus behaved as an electronic semi-conductor. The specific heat of  $\text{AgO}$ , as determined by Jirsa (1949), is given by  $0.0869 \pm 0.0005$ .

Specimens of the oxide,  $\text{AgO}$ , prepared by different methods,—such as the oxidation of metallic silver with ozone or a mixture of ozone and oxygen, oxidation of  $\text{Ag}^+$  with persulphate, the hydrolysis or thermal decomposition of  $\text{Ag}_7\text{NO}_{11}$ ,—showed the same powder patterns, and effectiveness as catalysts for the decomposition of ozone (Schwab and Hartman 1955).

Samples of higher oxides of silver were prepared by Graff and Stadelmaier (1958) by several chemical and electrochemical methods; viz. by the reaction between potassium persulphate and  $\text{AgNO}_3$  solution, by anodic oxidation of  $\text{AgClO}_4$  in aqueous perchloric acid or of  $\text{AgF}$  in aqueous hydrofluoric acid and of  $\text{AgNO}_3$  in aqueous nitric acid, as well as of porous sintered silver in aqueous caustic potash. Two distinct higher oxides of silver were obtained as products. The one with bipovalent silver,  $\text{Ag}_2\text{O}$ , was found to form a black powder and possess a monoclinic structure with lattice constants:  $a=5.842$ ,  $b=3.480$ ,  $c=5.487$  A.U.;  $\beta=107^\circ 27'$ . The authors found no evidence for the existence of other crystalline forms of  $\text{AgO}$ . Another higher oxide, also of black colour, was found to be associated in some proportions with  $\text{AgO}$  and to show a positive valency of greater than two with a face-centred cubic lattice. This higher oxide suffered slow decomposition to  $\text{AgO}$  on keeping for months.

The peroxide obtained by the anodic oxidation of silver salts is quite stable at room temperature, but suffers decomposition at higher temperatures, even in alkaline solution (Dirkse and Wiers 1959).

The crystal structure of the oxide,  $\text{AgO}$ , has been investigated by Scatturin *et al.* (1958) in order to determine the type of coordination and the bond distances of  $\text{Ag(II)}$  by the powder method of X-ray analysis. It has been found that the compound is monoclinic in structure and resembles very closely the structure of  $\text{CuO}$  (tenorite) with a planar configuration of

the oxygen atoms around the silver atom. The presence of two different Ag-O distances in the structure, viz. (1) 2.04 A.U. and (2) 2.35 A.U., which are commonly found in compounds of silver with formal valence of one and two respectively, has been discussed by the authors. Of the  $dsp^2$  hybridization, which should characterize Ag(II), only the planarity of the silver-oxygen groups is maintained.

The magnetic susceptibility of AgO has been measured by various workers with a view to obtain information regarding the valency state of the silver atom in the compound. A bipoisitive silver ion,  $Ag^{2+}$ , having one unpaired electron in the 4d shell should behave magnetically like  $Cu^{2+}$  ion with one unpaired 3d electron and a paramagnetic moment of about 1.73 B.M. But the oxide, AgO, is practically diamagnetic (Klemm 1931; Sugden 1932). The compound has been found to remain diamagnetic at temperatures between  $-183^\circ$  and  $+95^\circ C$ ;  $\chi_g$  at room temperature  $= 0.155 \times 10^{-6}$  (Neiding and Kazarnovskii 1951). The observed diamagnetism has been attributed by these latter authors to the formation of metallic or covalent bonding between the silver atoms in the crystal lattice. In the light of McMillan's study (1960) on the structure of AgO by X-ray analysis the diamagnetism of the compound may be readily accounted for, as it may then be represented as a mixed oxide of the composition,  $Ag_2O.Ag_2O_3$ , with the tripositive silver being linked to four oxygen atoms in a planar configuration. Alternatively, the compound might be regarded as a peroxide of silver having the composition  $Ag_2O_2$ . But the balance of evidence, as discussed above, is against the formulation of the compound as a peroxide.

On the other hand, paramagnetism has been observed for the compound obtained by the action of ozone on strongly acidified ( $HNO_3$ ) silver nitrate solution. It gave a moment value corresponding very closely to that expected for  $Ag^{2+}$  (Noyes, Pitzer and Dunn 1935). A moment of 1.76 B.M., which is in perfect agreement with the theoretical value for  $Ag^{2+}$ , has been reported for the peroxynitrate,  $Ag_7NO_{11}$ , by Neiding and Kazarnovskii (1951).

McMillan (1954, 1960) has closely investigated the structure of AgO by X-ray analysis and has obtained the same results as those reported by Graff and Stadelmaier (1958), as well as by Scatturin *et al.* (1958). The substance was found to form a monoclinic lattice belonging to the space group  $C_2/c - C_{2h}$  with the positions Ag(4d), O(4e). The choice of such a space group involves the assumption of four bivalent silver ions per unit cell. The results also indicate that silver ions are located in a face-centred cubic lattice, which gives the separation of neighbouring silver ions a value of 3.28 A.U. These observations agree equally well with the assumptions that all the four silver atoms are equivalent and occur as Ag(II), as well as that two of the silver atoms occur as Ag(III) and two as Ag(I) arranged in suitable planes with a somewhat lower symmetry space group  $P_{2_1}/c - C_{2h}$ . As AgO has been found to be diamagnetic, while a true silver(II) oxide

should show a paramagnetic moment corresponding to one unpaired electron, the formation of a metallic or a covalent Ag–Ag bonding has been postulated to account for this anomalous behaviour (Neiding and Kazarnovskii 1951). Such an explanation according to the author (McMillan 1960) is unsatisfactory since the metallic bonding should lead to much higher values of the conductivity, and the Ag(II)–Ag(II) covalent bonding between two bipoisitive kernels of small radius is not energetically feasible. He, therefore, assumes that in a unit cell of AgO, two of the silver atoms occur as Ag(III) and the other two as Ag(I). The tripositive silver ions are each coordinated with four oxygen, two of them being closer (dp hybridization) than the other two (sp hybridization). Of the four oxygen of the unit cell, two are the nearest neighbours of each Ag(I) ion, which suggests a linear sp coordination as is usual in covalent silver(I) compounds. The dark black colour of AgO has been attributed to the Ag–O bonding, the colour deepening with the increasing covalent character of the Ag–O bond (Helmholz and Levine 1942). The following distances between the neighbouring ions in AgO have been derived by the author (McMillan 1960).

Ag(III) – O  $\cong$  Ag(I) – O  $\cong$  2.1 A.U. ; this is comparable to the value found in Ag<sub>2</sub>O for the distance Ag(I) – O = 2.06 A.U.

O – O  $\cong$  2.8 A.U. > 2.64 A.U. (ionic diameter of oxygen)

Ag(III) – Ag(III) = Ag(I) – Ag(I) = 3.28 A.U.

Ag(III) – Ag(I) = 3.39 A.U. Cf. Fig. 1.

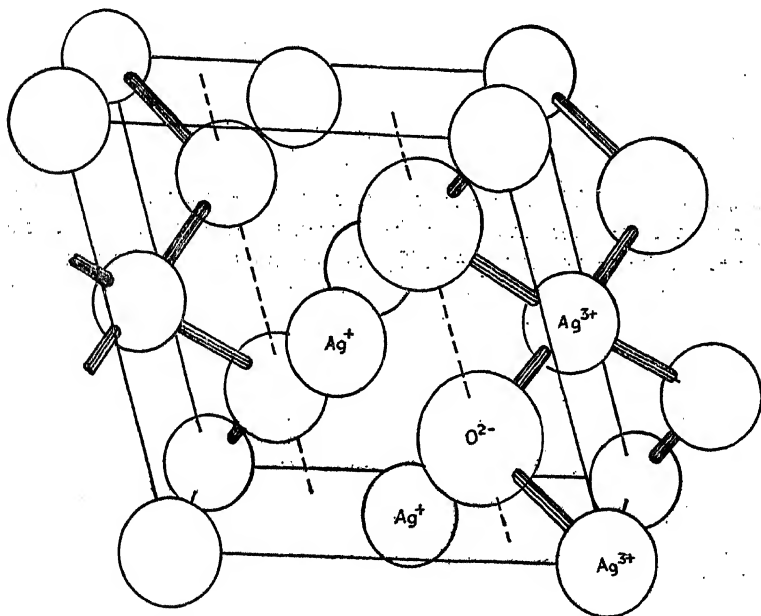


Fig. 1. Arrangement of atoms in AgO. (By courtesy, the author and Pergamon Press Ltd.)

It is claimed that most of the physical and chemical properties of AgO can be satisfactorily explained on this basis. This has been confirmed by the same author in his recent examination (private communication) of AgO in KOH by the method of paramagnetic resonance. Contrary to Dirkse and Wiers' (1959) finding he, therefore, concludes that there are no  $\text{Ag}^{2+}$  in alkaline solution; it breaks up into  $\text{Ag}^+$  and  $\text{Ag}^{3+}$  with possible formation of  $\text{Ag}(\text{OH})_2^-$  and  $\text{Ag}(\text{OH})_4^-$  respectively.

McMillan (private communication) has also studied the paramagnetic resonance spectrum of  $\text{Ag}^{3+}$  in frozen acid solutions ( $\text{HNO}_3$ ,  $\text{H}_2\text{SO}_4$ ) at  $77^\circ\text{K}$ , which can be interpreted as a doublet (interaction with silver nuclei) of strong g-anisotropy.

From all available evidences it appears now that in its combination with oxygen silver does not occur in the bipositive state, but either in the unipositive state as  $\text{Ag}_2\text{O}$  or in a mixed uni- and tripositive states as a molecular compound represented by the formula,  $\text{Ag}_2\text{O}.\text{Ag}_2\text{O}_3$ . The tripositive silver oxide,  $\text{Ag}_2\text{O}_3$ , has not been isolated in the free and pure state as yet.

An excellent review article by Dirkse (1959) on silver oxide electrodes and the oxides of silver has recently appeared in the Journal of Electrochemical Society to which a special reference might be made.

AgO has been found to serve as a powerful oxidimetric reagent in acid medium for analytical purposes. Methods for the volumetric estimation of vanadium, chromium, manganese and cerium have been described, using AgO as the oxidizing agent (Tanaka 1954; Lingane and Davis 1956). Colorimetric determination of vanadium has also been reported (Tanaka 1954).

#### (b) SILVER(II) FLUORIDE.

Silver(II) fluoride,  $\text{AgF}_2$ , has been prepared as a dark-brown or black solid by the direct combination of the elements at  $150^\circ\text{--}300^\circ\text{C}$  and by the action of fluorine on silver(I) halides (Ebert *et al.* 1933; Ruff and Giese 1934). The compound has also been obtained in the form of a yellow solid by heating powdered silver in a current of fluorine at  $250^\circ\text{C}$  in a magnesium vessel (Jockush 1934). Rochow and Kukin (1952) used chlorine trifluoride in the preparation of silver(II) fluoride from  $\text{AgCl}$ .

Silver(II) fluoride reacts violently with water yielding oxygen and ozone with formation of  $\text{Ag}_2\text{O}$  and an oxidizing solution which possibly contains fluoric acid. Silver fluoride is a powerful oxidizing as well as a fluorinating agent. It can easily oxidize ethanol to acetaldehyde, manganous salts to permanganate, chromic salts to chromates, etc. It reacts with  $\text{SO}_2$  and CO to form  $\text{SO}_2\text{F}_2$  and  $\text{COF}_2$ , and converts  $\text{CCl}_4$  to  $\text{CF}_4$  (Ebert *et al.* 1933; Jockush 1934).

The heat of formation of  $\text{AgF}_2$  by direct synthesis was found to be  $84.5 \pm 1.2$  kcal/mole, and its thermal stability is demonstrated by the fact that it melts at about  $690^\circ\text{C}$  with virtually no decomposition and that its decomposition pressure at  $700^\circ\text{C}$  is less than 0.1 atm. (Wartenberg 1939).

Silver(II) fluoride presents an interesting phenomenon in its magnetic behaviour. Gruner and Klemm (1937) obtained for  $\text{AgF}_2$  a surprisingly low molar susceptibility of  $+440 \times 10^{-6}$  at  $20^\circ\text{C}$  independent of the field. This is far below the theoretical value for one unpaired electron. But a sharp rise in the paramagnetic susceptibility occurs with the lowering of temperature when a weak ferromagnetism is attained at very low temperatures with a well-defined Curie point at  $-110^\circ\text{C}$ . The mass susceptibility at this point is given by the value  $270 \times 10^{-6}$ . It is, therefore, assumed that the compound shows antiferromagnetism with metallic (antiparallel spins) Ag-Ag bonds at high temperatures, but assumes a weak ferromagnetic character with parallel spins at low temperatures.

## II. COMPLEX COMPOUNDS.

The higher oxidation states of silver are usually stabilized by complex formation, particularly in combination with ligands offering the possibility of formation of dative  $\pi$ -bonds (Orgel 1956).

Many well-defined and crystalline complex compounds of both bi- and tripositive silver have been described in the literature. With few exceptions in the case of silver(II) complexes they all form 4-covalent penetration or inner level square planar complexes with  $\text{dsp}^2$  hybrid bonds (Rây 1946). All the bipositive silver complexes show a paramagnetic moment of c. 1.73 B.M. while the tripositive complexes are diamagnetic, in conformity with their electronic configurations as represented below.

	4d	5s	5p	5d
$\text{Ag}^{2+}$ (ion)	<u>1</u> <u>1</u> <u>1</u> <u>1</u> <u>1</u> <u>1</u> <u>1</u> <u>1</u>			
$\text{Ag(II)-complex}$ (4-coordinated)	<u>1</u> <u>1</u> <u>1</u> <u>1</u> <u>1</u> <u>1</u>			<u>1</u>
$\text{Ag(II)-complex}$ (6-coordinated)	<u>1</u> <u>1</u> <u>1</u> <u>1</u> <u>1</u> <u>1</u> <u>1</u> <u>1</u>			<u>1</u> <u>1</u>
$\text{Ag}^{3+}$ (ion)	<u>1</u> <u>1</u> <u>1</u> <u>1</u> <u>1</u> <u>1</u> <u>1</u>			
$\text{Ag(III)-complex}$ (4-coordinated)	<u>1</u> <u>1</u> <u>1</u> <u>1</u> <u>1</u> <u>1</u>			
Associated or outer level complexes with $\text{sp}^3$ or $\text{sp}^2\text{d}$ bonds.				
$\text{Ag(II)-complex}$ (4-coordinated)	<u>1</u> <u>1</u> <u>1</u> <u>1</u> <u>1</u> <u>1</u> <u>1</u> <u>1</u>			

Squares covered by the brackets indicate orbitals available for bond formation.

But the properties of the bipositive silver complexes with moment values of c. 1.73 B.M. with practically little contribution from the orbital moment justify their representation as inner level complexes with  $dsp^2$  hybrid bonds. With the unpaired d-electron promoted to the outermost level, exposed to the influence of the field of the neighbouring atoms and ions, there is every likelihood of its orbital moment being completely quenched (Rây and Sen 1948).

It is also worthy of note that the number of 6-coordinated or octahedral complexes of bipositive silver is rather limited. Even the tripositive silver with its higher charge has not yet been found to give any 6-coordinated complex. This might be attributed to the comparatively high electronegativity of silver which in consequence derives an adequate share of electrons from a relatively smaller number of ligands to satisfy Pauling's "postulate of neutrality" (Pauling 1945). In this respect tripositive silver resembles closely the tripositive gold.

#### (a) SILVER(II) COMPLEXES.

The bipositive state of silver has been found to be stabilized in complexes with a number of organic ligands of the type of heterocyclic nitrogen bases. The oxidation potential of Ag(I) to Ag(II) is rather high, being 1.914 volts (Noyes and Kossiakoff 1935) in acid ( $HNO_3$ ) solution and 1.41 volts (Luther and Pokorny 1908) in alkaline solution; but these values are considerably lowered by coordination. This has made possible the preparation of a number of more or less stable bipositive silver complexes with suitable ligands, particularly of the chelating type.

The orange-coloured prismatic crystals of  $[Ag^{II}(py)_4]S_2O_8$  (where  $py$  = pyridine), isomorphous with  $[Cu^{II}(py)_4]S_2O_8$ , were obtained by Barbieri (1912) by oxidizing an aqueous solution of silver nitrate in pyridine with potassium persulphate. The complex persulphate is unaffected by light. The corresponding nitrate,  $[Ag^{II}(py)_4](NO_3)_2$ , was prepared electrolytically in the form of orange-red needles from aqueous pyridine solution in a divided cell. The compound was isolated in small yield from the anode liquor which was composed of 8-10 per cent. silver nitrate and 40 per cent. pyridine, the cathode solution being 10-15 per cent. sulphuric acid placed in a porous cup. A current of 0.2-0.6 amp. at 5-6 volts between Pt-electrodes was employed for the preparation.

Mixed crystals of silver(II) and cadmium(II) pyridine persulphate,  $[Ag^{II}Cd^{II}(py)_4]S_2O_8$ , have been obtained by oxidizing a mixture of silver(I) and cadmium(II) salts with  $K_2S_2O_8$  in presence of pyridine. This serves as an evidence that silver occurs in the bipositive state in the silver pyridine complex prepared by oxidation (Barbieri 1931).



Magnetic properties of mixed crystals of silver(II) and copper(II) pyridine complexes,  $[\text{Ag}_x^{\text{II}}\text{Cu}_{(1-x)}^{\text{II}}(\text{py})_4]\text{S}_2\text{O}_8$ , have been studied by Capatos and Perakis (1936). The Curie point and the Curie constant were both found to vary with concentration ( $x$ ). The curves show at first a linear drop in the constant and a rise in the point as  $x$  changes from either 0 or 1, followed in both cases by an abrupt reversal in direction to a maximum for the constant and a minimum for the point at  $x = 0.5$  (Fig. 2).

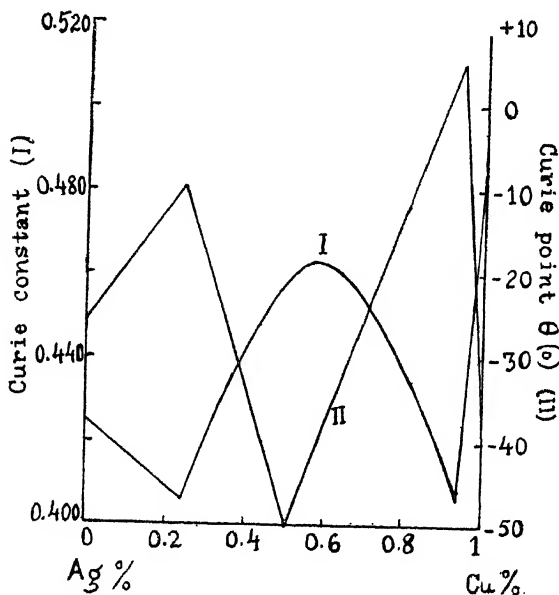


Fig. 2

Perakis and Capatos (1938, 1939) have made a thermomagnetic study of the complexes  $[\text{Ag}_x^{\text{II}}\text{Cd}_{(1-x)}^{\text{II}}(\text{py})_4]\text{S}_2\text{O}_8$  and  $[\text{Cu}_x^{\text{II}}\text{Cd}_{(1-x)}^{\text{II}}(\text{py})_4]\text{S}_2\text{O}_8$  in the liquid air to room temperature interval for a wide range of concentrations of silver and copper. In mixtures very weak in copper and silver the curves giving the variation of the magnetic moment and Curie point as functions of the concentration have approximately the same form and direction in the two systems. In mixtures rich in copper and silver the moments vary but little with concentration. In the more dilute solutions at a certain concentration the moment increases suddenly and approaches, at least for silver, the theoretical moment of the  $\text{Ag}^{3+}$  ion (14 Weiss magnetons). Cf. Fig. 3. If it be assumed as an explanation that silver(II) suffers disproportionation into silver(III) and silver(I) on magnetic dilution, then there would occur some reduction in the net moment value of the system, assuming the complex to be ionic. The authors' explanation is therefore untenable. This

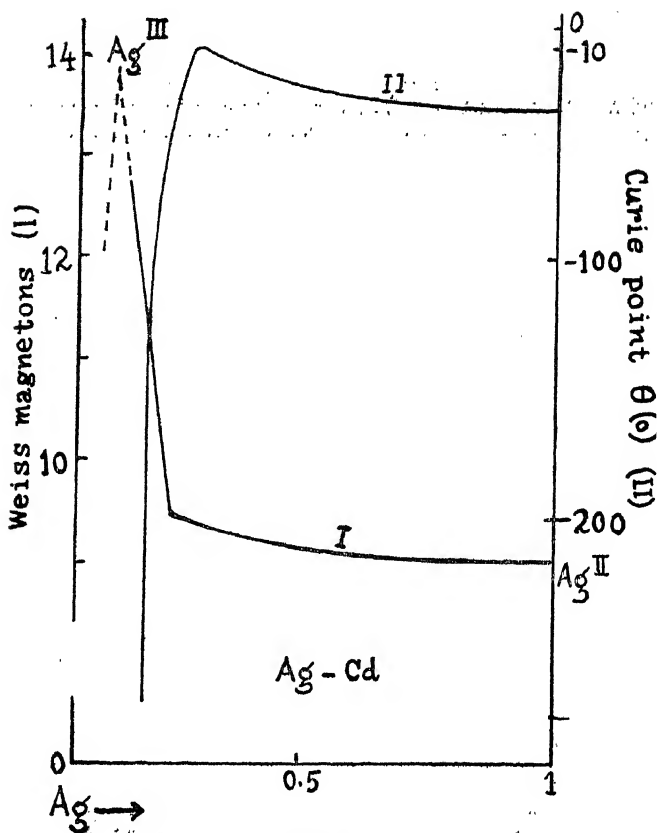
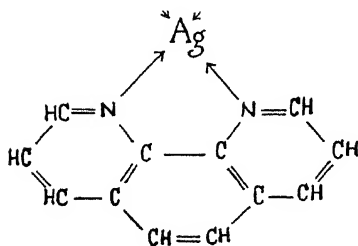
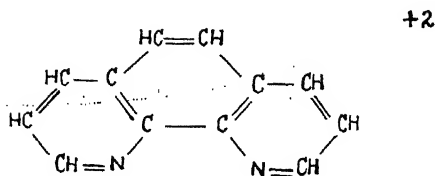


Fig. 3.

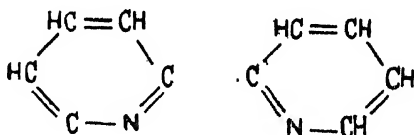
strange observation consequently arouses some suspicion, regarding the purity of the samples under study, or of their adventitious contamination with traces of paramagnetic or ferromagnetic impurities.

Hieber and Mühlbauer (1928) have described the preparation of reddish brown  $[\text{Ag}^{\text{II}}(\text{o-phen})_2]\text{S}_2\text{O}_8$  (o-phen = o-phenanthroline) in quantitative yield from a solution of silver nitrate by oxidation with ammonium persulphate in presence of an excess of o-phenanthroline. From a solution of the complex persulphate in 64 per cent.  $\text{HNO}_3$  the corresponding perchlorate, chlorate, nitrate and hydrogen-sulphate have been obtained by treatment with the appropriate alkali salts. The phenanthroline complexes are quite stable and are isomorphous with the corresponding copper(II) and cadmium salts. The complex silver(II) phenanthroline ion is represented by the configuration I.



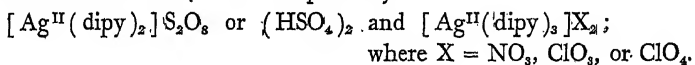
I

Using  $\alpha$ -dipyridyl (vide II) as the chelating ligand Morgan and Burstall (1930) prepared a number of complex silver(II) salts in the form of red



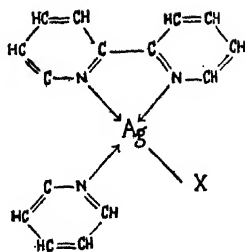
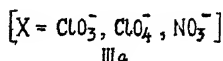
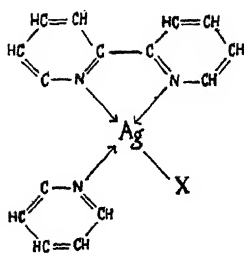
II

crystals from hot aqueous alcoholic solution of silver nitrate by oxidation with potassium persulphate. Two types of compounds were isolated by the authors with two and three molecular proportions of  $\alpha$ -dipyridyl showing coordination numbers of 4 and 6 respectively:



A binuclear compound of the composition,  $[\text{Ag}_2^{\text{II}}(\text{dipy})_5](\text{S}_2\text{O}_8)_2$ , has also been described. In a similar manner tripyridyl complexes of bipovalent silver of composition  $[\text{Ag}^{\text{I}}(\text{trpy})\text{X}]\text{X}$  and  $[\text{Ag}^{\text{II}}(\text{trpy})\text{S}_2\text{O}_8]$  have been prepared by the same authors (1937), and are represented by square planar structure (III a and b).

Nitrate, perchlorate and bisulphate of silver(II) dipyridyl complex,  $[\text{Ag}^{\text{II}}(\text{dipy})_2]\text{X}_2$  (where X =  $\text{NO}_3$ ,  $\text{ClO}_4$  or  $\text{HSO}_4$ ), have also been prepared by electrolytic oxidation in a divided cell (Barbieri 1932).

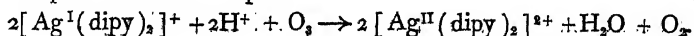


In the preparation of silver(II) complexes of  $\alpha\alpha$ -dipyridyl and *o*-phenanthroline, potassium persulphate can be replaced by the common oxidizing agents like  $\text{PbO}_2$ ,  $\text{BaO}_2$ ,  $\text{CeO}_2$ , etc., none of which are effective in oxidizing  $\text{Ag}$  to  $\text{Ag}^{2+}$  in the absence of any complexing agent (Malaguti 1950).

All the silver(II) complexes are strong oxidizing agents. In their molecules the silver atom is coordinated with 4 or 6 (in *tris*-dipyridyl complexes) nitrogen atoms provided by the ligand molecules, forming a square planar or octahedral structure as the case may be.

The preparation of a deep green silver complex of oxine, having the composition,  $[\text{Ag}(\text{oxine})_2]$ , was reported by Nakatsuka (1936), in which the silver atom was believed to occur in the bipovalent state. This was precipitated from a hot solution of silver(I) acetate by the addition of 8-hydroxyquinoline (oxine) without the use of any oxidizing agent. But Block, Bailar and Peace (1951) have shown that the compound, which can be obtained in two coloured forms, yellow and green, gives no evidence for any higher oxidation state of silver present in it. Moreover, the green variety, which was examined magnetically, was found to be diamagnetic like all compounds of univalent silver.

Malaguti and Labianca (1954) have made use of ozone as an oxidizing agent for the preparation of silver(II) complexes. In the case of dipyridyl salts,  $[\text{Ag}^{\text{II}}(\text{dipy})_2]\text{X}_2$  ( $\text{X} = \text{NO}_3$ ,  $\text{ClO}_3$  or  $\text{ClO}_4$ ), an almost quantitative yield has been reported. The following procedure was adopted for their preparation. The dipyridyl silver(I) nitrate,  $[\text{Ag}^{\text{I}}(\text{dipy})_2]\text{NO}_3$ , which separates as a precipitate on mixing a solution of silver nitrate with dipyridyl, was dissolved in the particular acid containing the anion of the complex salt under preparation. The solution was then saturated with the sodium salt of the anion involved and treated with ozone until the precipitation of the complex salt was complete.



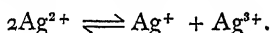
Ozone does not oxidize dipyriddy. The procedure is applicable to the preparation of silver(II) complexes of phenanthroline and other organic bases.

The standard oxidation-reduction potential of silver(II) dipyriddy complexes has been determined by Scrocco, Marmani and Mirone (1950). The value obtained is given by 1.453 volts. This is much smaller than those found for the system  $\text{Ag}^+ - \text{Ag}^{2+}$  in  $\text{HClO}_4$  and  $\text{HNO}_3$  solutions, viz. 2.0 and 1.929 volts respectively. The temperature coefficient also of the redox potential of the system,

$[\text{Ag}^{\text{I}}(\text{dipy})_2]^+ - [\text{Ag}^{\text{II}}(\text{dipy})_2]^{2+}$ , in nitric acid solution has been determined by Scrocco and Marmani (1954) and the value found is 0.0012 volt/degree.

Solutions of  $[\text{Ag}^{\text{II}}(\text{py})_4]^{2+}$ ,  $[\text{Ag}^{\text{II}}(o\text{-phen})_2]^{2+}$  and of  $[\text{Ag}(\text{oxine})_2](?)$  are all coloured red or brown with strong absorption at 3500–4000 A.U. (Sone 1949).

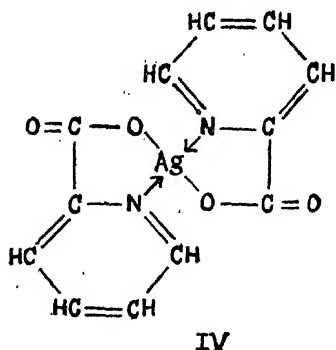
Bruno and Santoro (1956) studied the exchange reaction of the radioactive  $\text{Ag}^{110}$ -ion between silver nitrate and complex silver(II) compounds of  $\alpha$ -dipyriddy and *o*-phenanthroline. The silver(II) complex was separated by precipitation with  $\text{NaClO}_4$  solution and  $\text{Ag}^+$  by  $\text{HCl}$ . It was found that the total exchange occurred after 2 minutes. This seems to suggest that  $\text{Ag}^{2+}$  resulting from the dissociation of the complex suffers a partial disproportionation leading to the equilibrium:



In these silver(II) complexes, which are usually 4-covalent, the bonds are assumed on the basis of Pauling's theory (1945) to be of the hybrid  $\text{dsp}^2$  type giving rise to a square planar structure with a spare electron in the  $p_z$  orbital. The smaller difference in the free energy between silver(I) and silver(II) complexes of  $\alpha$ -dipyriddy and *o*-phenanthroline than the corresponding difference in their ionic states  $\text{Ag}^+$  and  $\text{Ag}^{2+}$  has been attributed by Scrocco and Ragazzini (1954) to the formation of covalent bonds with donor groups plus the stabilizing effect of the unpaired  $p_z$  electron resonating with similar electrons of the aromatic nuclei. With a view to confirm this assumption the authors made the measurement of redox potentials of silver(I) and silver(II) complexes of both  $\alpha$ -dipyriddy and *o*-phenanthroline. A decrease of 65 mv in the redox potential was observed in the case of phenanthroline complexes, since the number of electrons with which the silver(II)- $p_z$ -electron can interact is increased by 4 in the case of phenanthroline complex over that of the dipyriddy complex.

A very stable chelate complex of bipoisitive silver of the inner metallic type with picolinic acid (pyridine-2-carboxylic acid),  $[\text{Ag}^{\text{II}}(\text{C}_6\text{H}_4\text{NO}_2)_2]$ , was described by Barbieri (1933). It was obtained in the form of sparingly soluble, orange-red needles by the oxidation of silver nitrate solution with potassium persulphate in the presence of an excess of picolinic acid, rendered

alkaline with sodium carbonate. Electrolytic oxidation with platinum electrodes was also found equally effective for its preparation. In the presence of a copper salt mixed crystals containing both silver and copper were found to separate. The complex is decomposed by HCl with evolution of  $\text{Cl}_2$  and precipitation of AgCl. Ammonia and hydrogen peroxide are oxidized by it with evolution of  $\text{N}_2$  and  $\text{O}_2$  respectively. Cox *et al.* (1936) have shown by X-ray analysis that the compound has a square planar configuration IV and is isomorphous with coplanar copper complex of the same ligand. It shows moreover the high birefringence expected for a parallel arrangement of coplanar molecules.

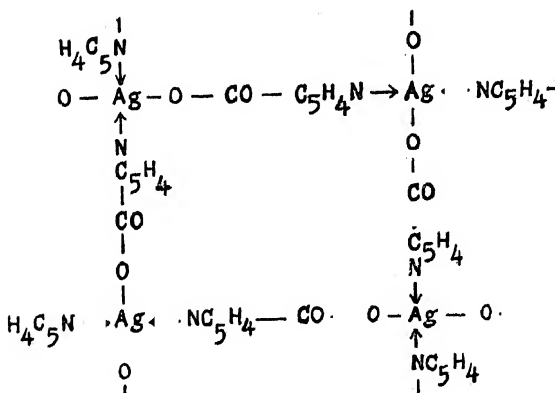
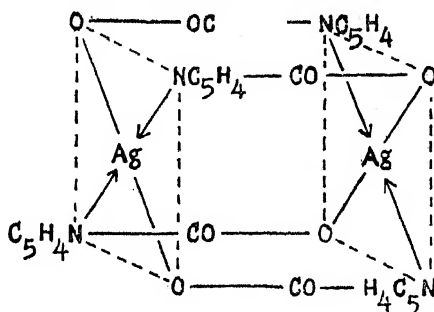


Silver(II) picolinate is, therefore, a quadricovalent complex with  $\text{dsp}^2$  hybrid bonds.

Banerjee and Rây (1956) have described the silver(II) complexes of nicotinic acid (pyridine-3-carboxylic acid) and of isonicotinic acid (pyridine-4-carboxylic acid). These were obtained in the form of insoluble cinnamon-red crystals by treating a cold aqueous solution of the chelating ligand concerned with silver nitrate and potassium persulphate solution. The complexes are more or less readily decomposed by dilute mineral acids and alkalis. They show a paramagnetic moment of c. 1.73 B.M., corresponding to the presence of one unpaired electron expected for a simple or complex bipoisitive silver compound. The silver(II) isonicotinate was found to be isomorphous with the corresponding copper(II) compound from an X-ray study of their powdered crystals. An indication of similar isomorphism for the nicotinate complexes of bipoisitive silver and copper has also been reported (Chackraburttty and Banerjee 1955) in view of the general resemblance of their X-ray patterns. From a consideration of their properties they are best represented as inner metallic quadricovalent silver(II) complexes with coplanar  $\text{dsp}^2$  hybrid bonds. But because of the possibility of steric hindrance to the bridging of meta- and parapositions of the aromatic ligands with formation of six- and seven-membered rings, the authors re-

present their structure as made up of either dimeric molecules as in V, or of polymeric molecules as in VI with the formation of an unending two-dimensional sheets and layers. An X-ray analysis of their structures is needed to arrive at a final conclusion on the matter.

Silver(II) complexes of the five out of six different pyridine dicarboxylic acids, viz., quinolinic (2:3), cinchomeronic (3:4), isocinchomeronic (2:5),



VI

lutidinic (2:4) and dipicolinic (2:6) acids have also been described by Banerjee and Rây (1957). These were prepared in the same manner as the silver(II) complexes of pyridine monocarboxylic acids and form more or less sparingly soluble red to dark-red crystals. The silver(II) dipicolinate was isolated, however, in two differently coloured modifications, dark-red (with  $2\text{H}_2\text{O}$ ) and green (with  $4\text{H}_2\text{O}$ ), differing in their water of hydration. The composition of these silver(II) complexes of pyridine dicarboxylic acids is represented by the general formula  $[\text{Ag}^{\text{II}}(\text{BH})_2]$ , where  $\text{BH}_2$  = one molecule of the acid concerned. Obviously one of the two  $-\text{COOH}$  groups

of each ligand molecule remains free and inactive. All these complexes are more or less readily decomposed by dilute mineral acids and alkalis. They all showed, as expected, a magnetic moment of 1.73 B.M. approx. The X-ray diffraction patterns with data for the silver(II) quinolate and for the two (dark-red and green) modifications of silver(II) dipicolinate have also been furnished by the authors. The preparation of a silver(II) quinolate with  $\text{H}_2\text{O}$ ,  $[\text{Ag}^{\text{II}}(\text{BH})_2] \cdot \text{H}_2\text{O}$ , which thus differs from that described by Banerjee and Rây in its water of hydration, was reported much earlier by Burada (1935). This was, however, prepared by oxidizing *bis*-quinoline silver(I) nitrate,  $[\text{Ag}^{\text{I}}(\text{C}_9\text{H}_7\text{N})_2]\text{NO}_3$ , with ammonium persulphate in aqueous solution.

The structure of silver(II) complex of cinchomeric acid is presumably related to that of nicotinic and of isonicotinic acid, being either dimeric, or polymeric with the formation of unending two-dimensional sheets or layers. Silver(II) complexes of all the other pyridine dicarboxylic acids are probably represented by a monomeric structure like that of silver(II) picolinate (Banerjee and Rây 1957).

Banerjee and Rây (1957) have further reported the preparation of silver(II) complexes of some of the pyridine tricarboxylic acids, namely, pyridine-2:3:6-, -2:4:5- and -2:4:6-carboxylic acids by oxidation of silver nitrate in cold aqueous solution with persulphate in presence of the acid concerned. The silver(II) complex of 2:3:6-tricarboxylic acid of pyridine, however, forms a very unstable green compound, decomposing even at  $5^\circ\text{C}$  in vacuum. Pyridine-2:4:6-tricarboxylic acid (collidinic acid) like dipicolinic acid gave two differently coloured products, chocolate and black. The former was found to crystallize with one molecule of water and the latter to separate in the anhydrous state. The silver(II) complex of pyridine-2:4:5-tricarboxylic acid forms beautiful orange-coloured microcrystals. The properties of these compounds are more or less similar to those of the silver(II) complexes of mono- and dicarboxylic acids of pyridine. They show a magnetic moment value of about 1.73 – 1.82 B.M. Their compositions can be represented by the general formula,  $[\text{Ag}^{\text{II}}(\text{BH}_2)_2]$ , where  $\text{BH}_2$  = one molecule of the pyridine tricarboxylic acid concerned. Two of the three  $-\text{COOH}$  groups in each molecule of the ligand remain free. The X-ray powder photographs with data for the two forms of silver(II) collidinate, chocolate and black, have also been presented by the authors.

The stability of the silver(II) complexes of pyridine mono-, di- and tricarboxylic acids has been found to vary in the order complex (monoacid) > complex (diacid) > complex (triacid). Within each group again the stability varies depending on the relative positions of the functional groups and of the inactive  $-\text{COOH}$  groups where present. The steric effect that hinders the formation of favourable planar configuration around the central silver atom reduces the stability of the complex.



All the quadricovalent silver(II) complexes, that have been examined magnetically, show a paramagnetic moment value of 1.73 B.M. approx. in perfect agreement with that expected theoretically for one unpaired electron (Klemm 1931; Morgan and Sugden 1931; Sugden 1932; Capatos and Perakis 1936; Banerjee and Rây 1956, 1957). A few silver(II) dipyriddy complexes described by Morgan and Burstall (1930), which show a coordination number of six, give on the other hand a somewhat higher moment value of 2.03 – 2.16 B.M. In these latter, therefore, the orbital moment is not completely quenched. For, a sexicovalent silver(II) complex, like the corresponding copper(II) complex, can at best be an associated or outer level complex with  $sp^3d^2$  hybrid bonds (Rây and Sen 1948).

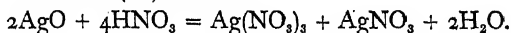
Paramagnetic resonance studies on the *bis-o*-phenanthroline silver(II) persulphate,  $[Ag^{II}(o\text{-phen})_2]S_2O_8$ , indicate that the electron hole in the 4d subshell is not localized on the silver(II) ion, but migrates to the surrounding atoms (Bowers 1953). This seems to suggest the formation of  $\pi$ -bonding.

The silver(II) dipyriddy complex,  $[Ag^{II}(dipy)_2]^{2+}$  has been found to serve as an oxidimetric titrant for the determination of oxalic acid, nitrous acid and quadrivalent vanadium (Barbieri and Labianca 1954).

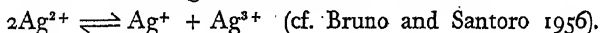
#### (b) SILVER(III) COMPLEXES.

There are evidences for the occurrence of  $Ag^{3+}$  in the form of simple salts in solution, though it has not been possible to isolate them in a pure solid state.

Barbieri (1931) studied the product formed by dissolving AgO in conc.  $HNO_3$ . The ratio of Ag:O (active) in the solution was determined by oxidimetric titration with various reagents. A ratio of 1:1.5-1.6 was obtained in all cases indicating the existence of  $Ag^{3+}$  in solution. The author, therefore, suggested that AgO dissolves in nitric acid with disproportionation into silver(I) and silver(III) nitrates:



From a study of the kinetics of the isotopic exchange reaction between Ag(I) and Ag(II) in perchloric acid solution Gordon and Wahl (1958) have proposed that the exchange occurs via the reaction,



$Ag^{3+}$  has, however, been remarkably stabilized by complex formation with some suitable ligands capable of yielding dative  $\pi$ -bonds and/or polycyclic rings with the metal atom.

The electrolytic oxidation of telluric acid at a silver anode is said to yield a compound having the composition represented by the formula,  $5Ag_2O_3 \cdot Ag_2O \cdot 9TeO_3$  (Brauner and Kuzma 1907). No analytical results have, however, been furnished by the authors. Apparently it is a basic salt containing some univalent silver.

Malaprade (1940) reports the isolation of a crystalline compound from the brown solution resulting from the oxidation of  $\text{Ag}_2\text{O}$  with a boiling concentrated solution of  $\text{KOH}$  and  $\text{KIO}_4$ . The product after recrystallization from aqueous  $\text{KOH}$  gave on analysis the atomic ratio of  $\text{K}:\text{Ag}:\text{I} = 6:1:2$ . On treating the brown solution with  $\text{NaOH}$  a salt was precipitated in which the ratio  $\text{Na}:\text{K}:\text{Ag}:\text{I}$  was found to be equal to  $4.5:1.5:1:2$ . The determination of available oxygen demonstrated the tripositive character of the silver atom in the compounds. With dilute acids the substances gave ozonized oxygen. The author has not given any definite composition for the products. They may obviously be represented by the formula,  $\text{M}_6\text{H}[\text{Ag}^{\text{III}}(\text{IO}_6)_2]$ ; where  $\text{M} = \text{K}$  or  $\text{K} + \text{Na}$ . A compound of this composition has subsequently been described by Malatesta (1941).

The preparation of a number of strongly coloured crystalline periodates of tripositive silver of the general formula,  $\text{M}_7[\text{Ag}^{\text{III}}(\text{IO}_6)_2] \cdot x\text{H}_2\text{O}$ , has been described by Malatesta (1941). The compounds are all diamagnetic and should for this reason, on the basis of valence bond theory, be regarded as quadricovalent complexes having square planar configuration with  $\text{dsp}^2$  hybrid bonds. In these the tripositive silver atom is covalently bound with four oxygen atoms, two from each periodate ( $\text{IO}_6$ ) group (cf. Rây and Sarma 1948). A tripositive silver atom, being isoelectronic with bipositive palladium, is likely to show a coordination number of four (vide Rây and Chakravarty 1944).

Oxidation of a mixture of  $\text{Ag}_2\text{SO}_4$ ,  $\text{KIO}_4$  and  $\text{KOH}$  with potassium persulphate in aqueous solution at the boiling temperature has led to the preparation of  $\text{K}_6\text{H}[\text{Ag}^{\text{III}}(\text{IO}_6)_2] \cdot 10\text{H}_2\text{O}$  in the form of orange crystals with a molar susceptibility ( $\chi_m$ ) of  $-40.5 \times 10^{-6}$ . By rapid evaporation of the reaction mixture a compound of the composition,  $\text{K}_7[\text{Ag}^{\text{III}}(\text{IO}_6)_2] \cdot \text{KOH} \cdot 8\text{H}_2\text{O}$ , has been obtained. Treatment of the former compound with dilute sodium carbonate solution yields the orange-yellow compound,  $\text{KNa}_6\text{H}[\text{Ag}^{\text{III}}(\text{IO}_6)_2] \cdot 16\text{H}_2\text{O}$ , with  $\chi_m = -31.6 \times 10^{-6}$ . This on treatment with dilute caustic soda solution and then with alcohol produces  $\text{KNa}_6[\text{Ag}^{\text{III}}(\text{IO}_6)_2] \cdot \text{NaOH} \cdot \text{H}_2\text{O}$  of a dirty yellow colour.

In a similar manner the same author (Malatesta 1941) has prepared some complex tellurates of tripositive silver of the general formula,  $\text{M}_7[\text{Ag}^{\text{III}}(\text{TeO}_6)_2] \cdot n\text{H}_2\text{O}$ . Oxidation with potassium persulphate of a mixture of  $\text{Ag}_2\text{SO}_4$ ,  $\text{TeO}_2$  and  $\text{NaOH}$  in water at the boiling temperature yielded the pale yellow crystals of the compounds,

$\text{Na}_6\text{H}_9[\text{Ag}^{\text{III}}(\text{TeO}_6)_2] \cdot 18\text{H}_2\text{O}$  with  $\chi_m = 22 \times 10^{-6}$ , and  $\text{Na}_7\text{H}_2[\text{Ag}^{\text{III}}(\text{TeO}_6)_2] \cdot 14\text{H}_2\text{O}$ . Like the corresponding periodates they should also be regarded as quadricovalent silver(III) complexes having square planar configuration with  $\text{dsp}^2$  hybrid bonds.

Malaprade and Coulombeau (1954) have found that the complex silver(III) periodate oxidized oxalic acid to  $\text{CO}_2$ . During the reaction, while  $\text{Ag}^{3+}$  is reduced to  $\text{Ag}^+$ , much of the periodic acid is simultaneously reduced to iodic acid. But as the cold dilute periodic acid has no action on oxalic acid, the reduction of periodic acid in the above reaction may be regarded as an instance of *induced reaction*, comparable to the reduction of mercuric chloride to mercurous chloride in a mixture of mercuric chloride and oxalic acid when the latter is oxidized with potassium permanganate in solution.  $\text{Ag}^{3+}$  behaves here in an analogous manner to permanganate, and periodic acid to mercuric chloride.

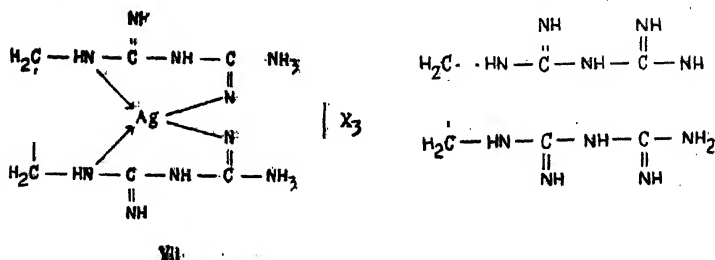
The preparation of a complex fluoride of tripositive silver,  $\text{KAgF}_4$ , has been reported by Klemm (1954). This is formed by the action of fluorine upon a mixture of  $\text{AgF}$  and  $\text{KF}$ . The substance was found to be diamagnetic.

Nakatsuka and Miwa (1952) have reported the preparation of a silver(III) complex with oxine by the action of silver(II) and silver(III) oxides on oxine in cold pyridine solution. It forms a yellowish brown, unstable product insoluble in water and many organic solvents, but highly soluble in pyridine. It oxidizes the halogen acids and decomposes in other acids.

Rây and Chakravarty (1944) have described a remarkably stable 4-covalent cationic complex of tripositive silver with ethylenedibiguanide. The complex base,  $[\text{Ag}^{\text{III}} \text{C}_2\text{H}_4(\text{C}_2\text{H}_4\text{N}_3)_2](\text{OH})_3 \cdot 3\text{H}_2\text{O}$ , is violet-red, while the salts, sulphate (with  $7\text{H}_2\text{O}$ ), nitrate, and perchlorate (with  $1.5\text{H}_2\text{O}$ ), form beautiful deep red needle-shaped crystals. The ethylenedibiguanide functions as a quadridentate chelating ligand and forms an inner metallic complex with the tripositive silver atom. The dibiguanide molecule behaves at the same time as a dibasic acid with two replaceable hydrogen atoms in two of its imino ( $=\text{NH}$ ) groups, as well as a diamine base with its two amino ( $-\text{NH}_2$ ) groups at the two ends. The complex base and its salts are all diamagnetic and are quite stable in the solid state. The nitrate can be recrystallized from hot dilute (2N) nitric acid without any change. As the electronic configuration of  $\text{Ag}^{3+}$  is closely analogous to that of  $\text{Pd}^{2+}$  and  $\text{Ni}^{2+}$ , the structure of the silver(III) ethylenedibiguanide compounds should, according to the authors, be represented by a square planar configuration with  $\text{dsp}^2$  hybrid bonds (VII).

The complex silver(III) ethylenedibiguanide sulphate was prepared by persulphate oxidation of a mixture of  $\text{Ag}_2\text{SO}_4$  and ethylenedibiguanide sulphate,  $\text{C}_2\text{H}_4(\text{C}_2\text{H}_4\text{N}_3)_2 \cdot 2\text{H}_2\text{SO}_4$  in cold aqueous solution. The substance is moderately soluble in water and loses its water of hydration ( $7\text{H}_2\text{O}$ ) at  $80^\circ\text{C}$ .  $\chi_g = -0.394 \times 10^{-6}$ . The complex base was obtained from the sulphate by treatment with cold  $\text{NaOH}$  solution. It is sparingly soluble in cold water and reacts strongly alkaline to litmus.  $\chi_g$  (base)  $= -0.39 \times 10^{-6}$ .

The complex nitrate and the perchlorate were prepared from the base by treatment with cold dilute (1*N*) nitric acid and 60% perchloric acid res-



pectively. The complex nitrate is moderately and the perchlorate is highly soluble in water.  $\chi_g$  (nitrate) =  $-0.402 \times 10^{-6}$ . The molecular conductivity of the nitrate, 518 ohm<sup>-1</sup> at 20°C, corresponds to that of a tripositive complex ion (Rây and Chakrabarty 1944). A microphotograph of the crystals of the complex nitrate is shown in Fig. 4a.

Attempts to prepare similar complexes of tripositive silver with the simple biguanide,  $C_2H_7N_5$ , ended in failure, as the ligand itself was oxidized by the persulphate under the conditions employed for the preparation of the silver(III) ethylenedibiguanide complex. Investigations of Noyes *et al.* (1935, 1937), as well as of Luther and Pokorny (1908), have shown that the oxidation potential of Ag(I)–Ag(II) increases with the  $p_H$  value of the solution. In alkaline solutions it has been found to be much lower than in acid solutions. This also holds good for the oxidation to Ag(III) state. It might, therefore, be expected that by a proper adjustment of the  $p_H$  value of the solution Ag(I) may be oxidized to a higher positive state by persulphate oxidation, leading to a complex formation with the simple biguanide without the ligand itself being oxidized at the same time. This has actually been realized recently by one of us. Sen (1960) has just succeeded in preparing the silver(III) biguanide complex of the composition,  $[Ag^{III}(C_2H_7N_5)_2](OH)SO_4 \cdot 7H_2O$ , in the form of sparingly soluble, brown silky crystals by oxidizing with potassium persulphate in cold aqueous solution a mixture of biguanide sulphate and silver(I) nitrate at a  $p_H$  6.5–7.0. The substance was found to be diamagnetic with  $\chi_g = -0.689 \times 10^{-6}$ . Fig. 4b shows a microphotograph of its crystals. The complex hydroxide-sulphate was converted to the orange-yellow nitrate sulphate,  $[Ag^{III}(C_2H_7N_5)_2](NO_3)SO_4 \cdot 6H_2O$ , by treatment with nitric acid (1*N*). From these the free complex base, as well as its nitrate and sulphate, have been obtained in the pure state. The complex is represented by a structure similar to that for silver(III) ethylenedibiguanide, preferably in a trans form. VIII. Similar complexes of tripositive silver with some alkylbiguanides have also been obtained in the form of orange-coloured crystals.



Fig. 4a. Ag(III) Ethylenedibiguanide nitrate crystals.

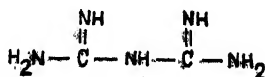
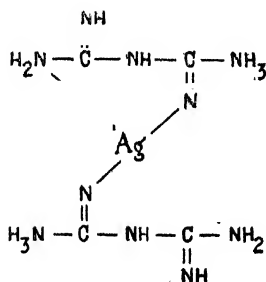


Fig. 4b. Ag(III) Biguanide hydroxide-sulphate crystals.



Curiously enough, with biguanide as a complexing ligand no bipositive silver complex has yet been obtained. An investigation with this end in view is, however, awaited with interest.

A solution of the complex silver(III) ethylenedibiguanide nitrate or sulphate has been found by the writers of the present review to show a wide band of absorbance maximum extending from 300–370  $m\mu$  with a molar extinction coefficient of 2325 (nitrate) and 2422 (sulphate) at 32°C.

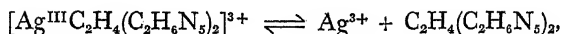


### VIII

Silver(II) complexes of pyridine and o-phenanthroline show, according to Sone (1949), a similar absorption band at 350–400  $m\mu$ . All these silver(II) and silver(III) complexes possess the same atomic configuration with the central silver atom being covalently linked to the four surrounding nitrogen atoms of the ligands in a square planar structure. The shifting of the absorption band towards the ultraviolet in the case of the silver(III) complex seems to suggest an increase in the strength of the Ag–N bond and hence an increased stability of the silver(III) complexes. This is not unexpected in view of the increased charge on the silver atom and is, in fact, supported by the general properties of the silver(II) and silver(III) complexes, as well as by the very high value of the instability constant of silver(III) ethylenedibiguanide nitrate in solution.

The stability of silver(III) ethylenedibiguanide nitrate in 2N KNO<sub>3</sub> solution has been determined by Sen, Ghosh and Rây (1950) from the equilibrium constant of the reaction between the complex salt and the hydrogen ion (HNO<sub>3</sub>), using the acid dissociation constants of the ethylenedibiguanide base. The equilibrium constant was derived from the p<sub>H</sub> and Ag<sup>3+</sup> concentration of the equilibrated solution. The Ag<sup>3+</sup> concentration was derived from a measurement of the redox potential of the equilibrated solution in the presence of Ag<sup>+</sup> ion. On the assumption of a single stage

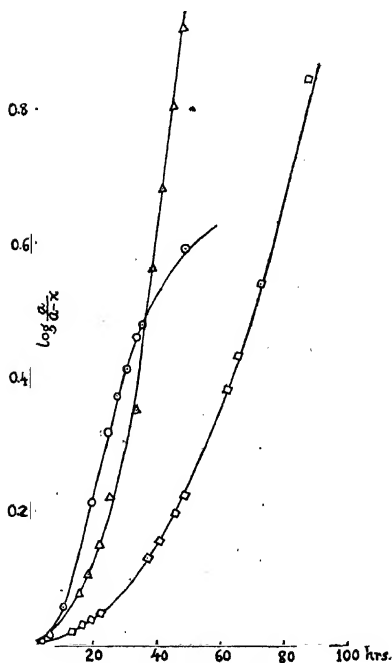
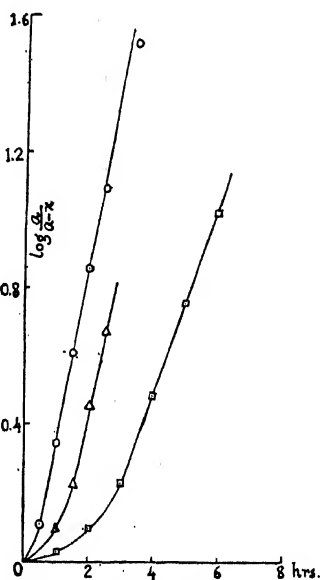
dissociation of the complex ion the value of its instability constant for the reaction ;



was found to be given by  $\text{pK} = 52.16$  at  $32^\circ\text{C}$ . The value was also corrected for the decomposition of  $\text{Ag}^{3+}$  ion in solution according to the equation,  $\text{Ag}^{3+} + \text{H}_2\text{O} \rightleftharpoons \text{Ag}^+ + 2\text{H}^+ + \frac{1}{2}\text{O}_2$ .

The rate of decomposition of silver(III) ethylenedibiguanide nitrate at different  $\text{pH}$  and temperatures ( $25^\circ$ ,  $35^\circ$  and  $45^\circ\text{C}$ ) has been studied by Sen and Rây (1953). The reaction rate is  $\text{pH}$  dependent, but the reaction proceeds in a rather complicated manner. The curves in Figs. 5 and 6 show the rate of decomposition of the complex nitrate in water and acid ( $\text{HNO}_3$ ) solutions at  $25^\circ$  and  $45^\circ\text{C}$ . It is apparent from the slope of the curves that

10

Fig. 5. temp. =  $25^\circ\text{C}$ .Fig. 6. temp. =  $45^\circ\text{C}$ .

—○—○— water solution.  
 —△—△— 0.025 M  $\text{HNO}_3$  solution  
 —□—□— 0.05 M  $\text{HNO}_3$  solution

$a$  = initial concentration of the complex

$(a-x)$  = concentration of the total tripositive silver after a time  $t$ .

the decomposition is somewhat retarded in acid solutions. It has been



suggested that in aqueous solution the decomposition occurs mainly through some intermediate stage with the formation of bipositive silver and hydrogen peroxide, while the direct decomposition of  $\text{Ag}^{3+}$  to  $\text{Ag}^+$  predominates in acid solutions. The mechanism of decomposition has been represented by the following equations:

In acid solutions:  $\text{Ag}^{3+} + \text{H}_2\text{O} \rightleftharpoons \text{Ag}^+ + 2\text{H}^+ + \frac{1}{2}\text{O}_2$ ;

In water:

$\text{Ag}^{3+} + \text{H}_2\text{O} \rightleftharpoons \text{AgO}^+ + 2\text{H}^+ \rightleftharpoons \text{Ag}^{2+} + \text{H}^+ + \frac{1}{2}\text{H}_2\text{O}_2 \rightleftharpoons \text{Ag}^+ + 2\text{H}^+ + \frac{1}{2}\text{O}_2$ .

The values of activation energy,  $E$ , for the decomposition reaction in the two cases have also been calculated and are given by 18,650 cal/mole in acid solution and 27,800 cal/mole in water solution.

In this monograph we have tried to cover references up to the latest publications on the subject, as far as they have been available to us. A glance at the list of references would indicate the wealth of materials accumulated on this topic of a rather limited character. We hope the present review will stimulate further interest in the subject, as there is still ample scope of intensive study on the chemistry of bi- and tripositive silver, particularly in respect of their complex compounds. For, such a study, in our opinion, is likely to reveal many new facts and ideas relating to valency and the structure of chemical compounds, besides elucidating the old ones.

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